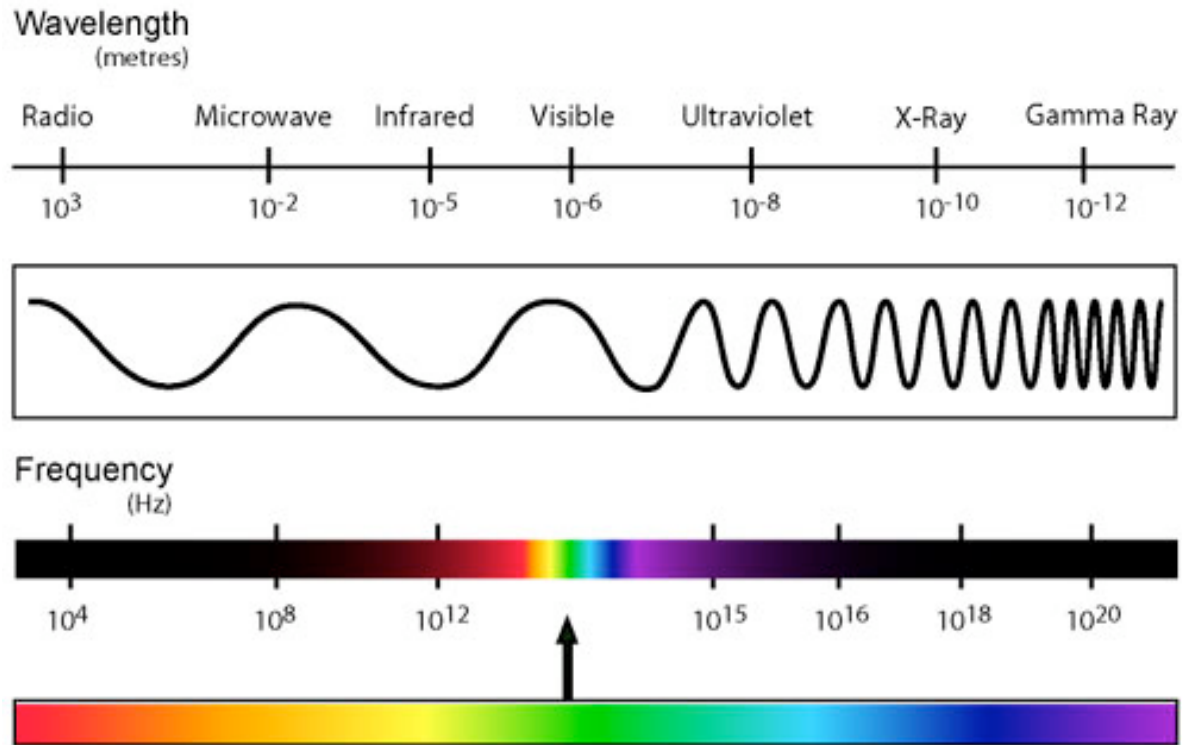


THE ELECTRO MAGNETIC SPECTRUM



$$c = \lambda \nu$$

c = speed of light
 3×10^8 m/s

ν = frequency
waves per second

λ = wavelength

Constructing electronic structure of a hydrogen atom

Part I:

energy of radiation can be estimated by its frequency

$$\nu = \frac{c}{\lambda}$$

ν : frequency of radiation

c : velocity of radiation

λ : wavelength of radiation

$\nu(\text{unit} = \text{s}^{-1})$

$\lambda(\text{unit} = \text{nm} ; 1 \text{ nm} = 1 \times 10^{-9} \text{ m})$

$c = 3 \times 10^8 \text{ m s}^{-1} / :300,000\text{km}$

$$E = h\nu$$

E : energy of radiation

h : Planck constant

ν : frequency of radiation

$E(\text{unit} = \text{J})$

$h = 6.626 \times 10^{-34} \text{ Js}$

Photoelectric effect:

The energy of an electron is proportional to the frequency of light given off.

The experiment:



Photoelectric Effect

Because each element has a different number of electrons, each element gives off characteristic wavelengths, or colors.

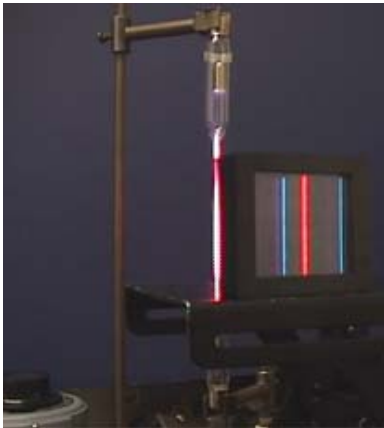


Flame Tests for Metals

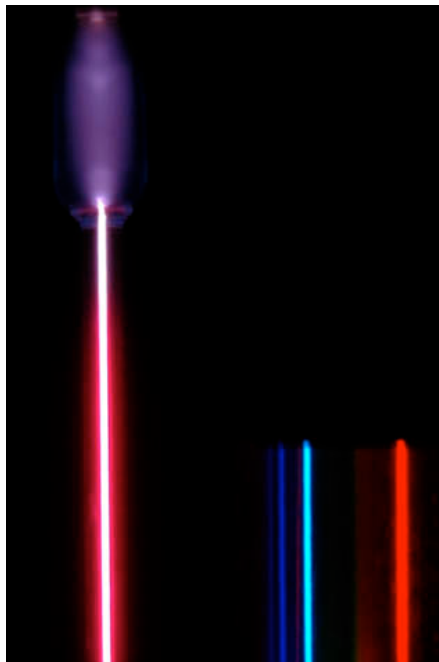
When electrons are excited, they increase in energy. When they drop down to their original location, they give off energy equal to the difference between the 2 locations.

So, the color of the flame tells us something about the energy levels the electrons occupied.

We can excite hydrogen atoms by running an electric current through hydrogen gas.....



And then viewing the glowing gas through a **spectroscope**



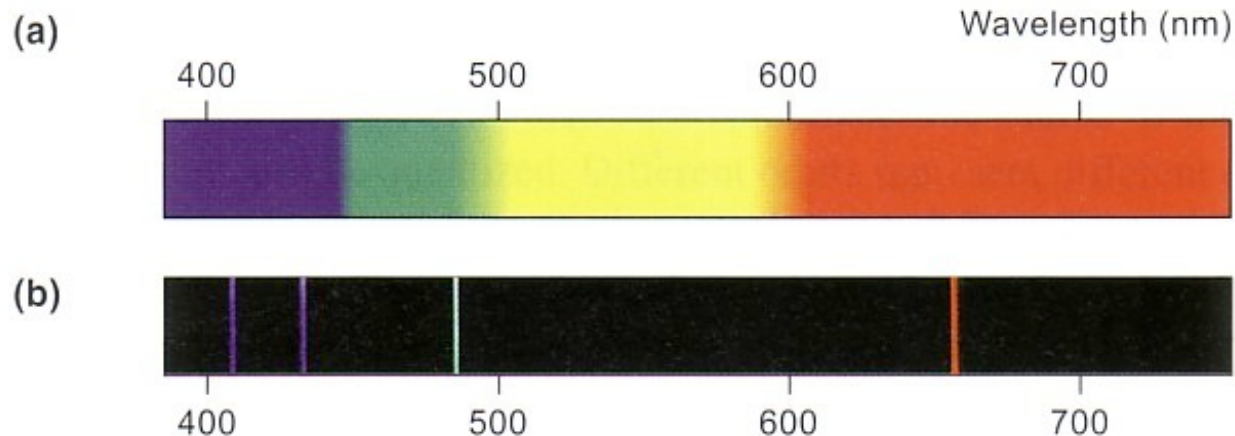
The result is the visible **hydrogen spectrum**:

This can give us a picture of the hydrogen atom:

Constructing electronic structure of a hydrogen atom

Part II:

Hydrogen atom emission spectrum give us hints ...



(a) is a *continuous spectrum* obtained from *visible light*, λ from 400nm to 700 nm.

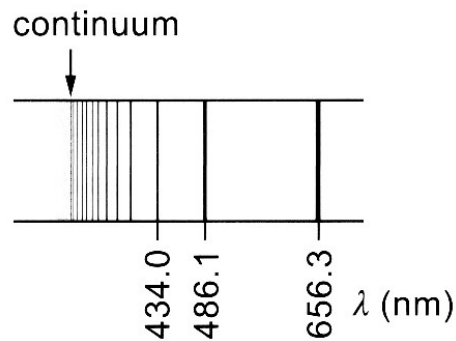
(b) is a *line spectrum* obtained from *emitted light* of an excited hydrogen atom, λ from 400 nm to 700 nm.

The H atom emits light at **discrete energy levels**.

Niels Bohr discovered that H's lone electron emits light in the discrete levels seen in the hydrogen spectrum:

Constructing electronic structure of a hydrogen atom

Bohr's enlightenment



Balmer series

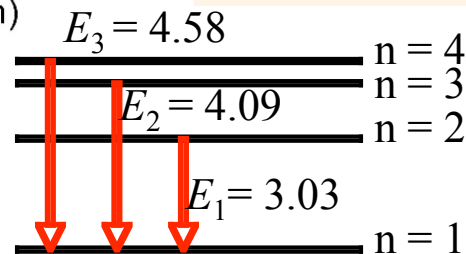
$$E_1 = 3.03 \times 10^{-19} \text{ J}$$

$$E_2 = 4.09 \times 10^{-19} \text{ J}$$

$$E_3 = 4.58 \times 10^{-19} \text{ J}$$



Niels Bohr (1885 -- 1962) und Max Planck in Kopenhagen, 1930



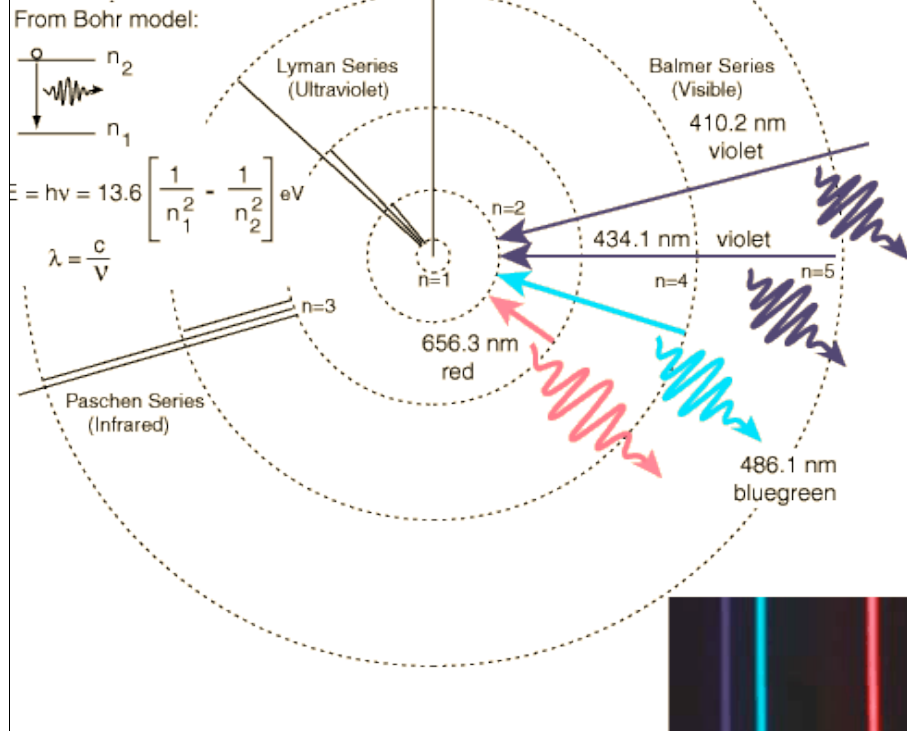
n = shell number

n is also called principal quantum number

Features of energy levels of H atom:

1. discrete / 2. energy gap decreasing

Bohr used this concept of **energy levels** to construct his model of the hydrogen atom:



Features of Bohr's model:

Electrons occupy discrete energy levels, or “shells”

The energy gap decreases as the shell number increases

So if hydrogen's spectrum can tell us about the structure of a H atom, can the spectra of other elements help us create models of those atoms?

<http://jersey.uoregon.edu/vlab/elements/Elements.html>

What do we notice about the spectra of the other elements?

They are more complex than H

In general, their complexity increases as atomic number increases

The Bohr Model **does not** work for elements other than Hydrogen.

Atomic spectra just get too complicated to be explained by the Bohr Model.

So, a new model must be created to explain the atomic structure of **all** the elements, not just hydrogen!

Enter Schrodinger used **probability** to express the location of an electron in an atom by developing a mathematical equation, the Schrodinger Equation.

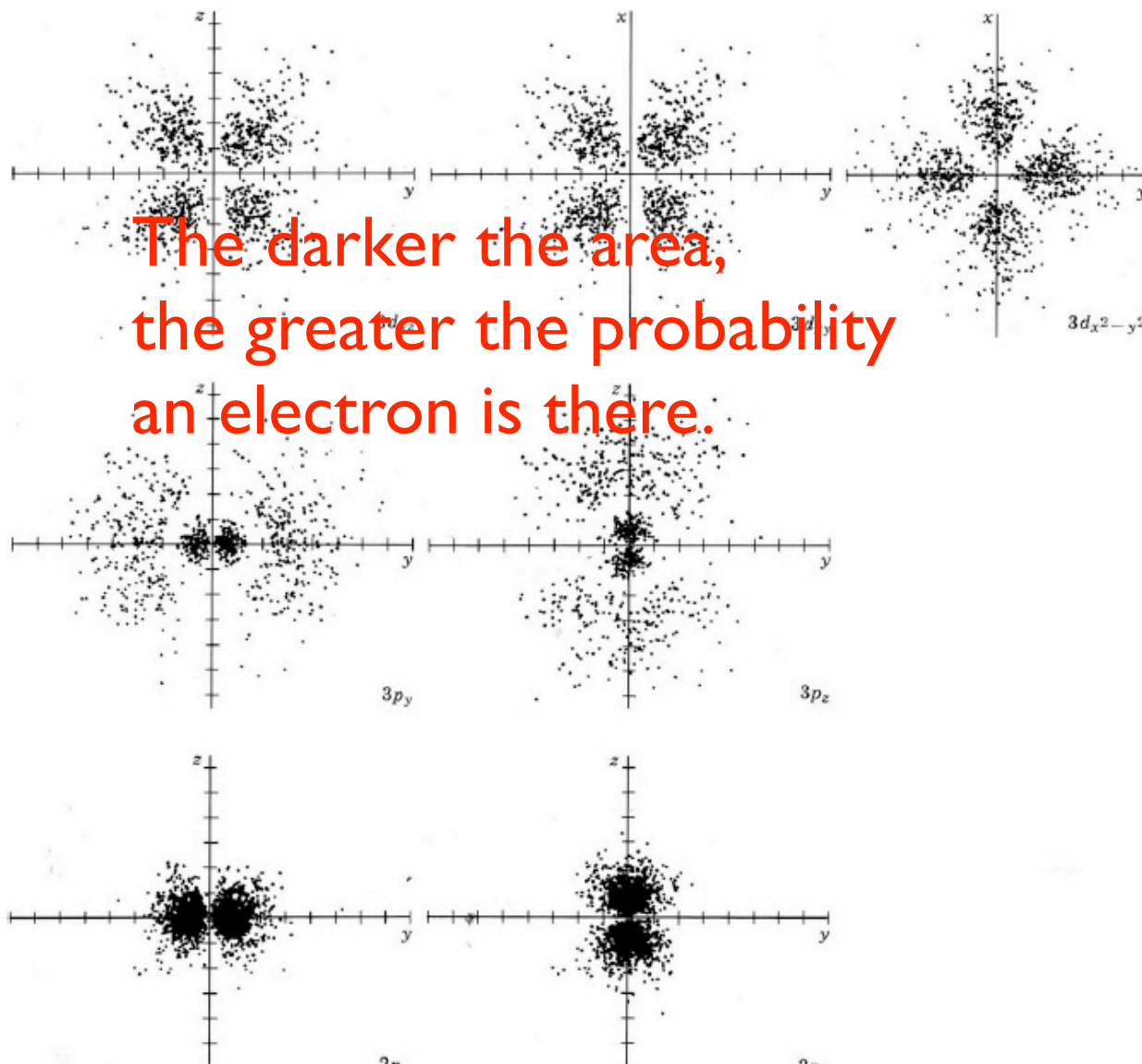


I developed the Schrodinger equation, and quantum mechanics!

This led to the development of the **quantum mechanical model** of the atom.

Let's compare the **Bohr model** with the **quantum mechanical model**.

Characteristics of the quantum mechanical model:



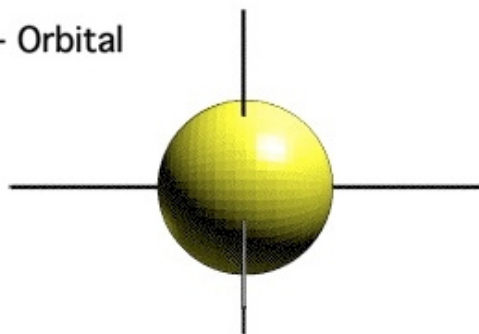
The darker the area,
the greater the probability
an electron is there.

The location of electrons are given by the probability of finding them in that area. These areas have different energy levels.

These probability regions are called **atomic orbitals**.

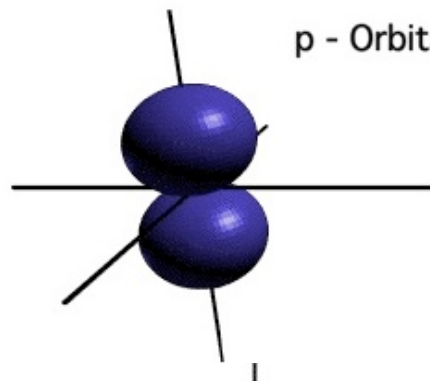
These orbitals have the following shapes:

s - Orbital



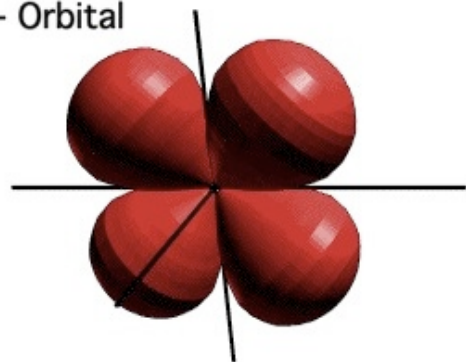
each can contain 2 electrons
Groups 1 and 2 on the periodic table

p - Orbital



each can contain 6 electrons
Groups 3-8 on the periodic table

d - Orbital



each can contain 10 electrons
transition metals on periodic table

Energy levels of electrons are denoted by principal quantum numbers (n). $n = 1, 2, 3, 4$, etc.

Each energy level, n , may have energy sublevels, each corresponding to an orbital of a different shape.

Principal
energy
level

Number
of sublevels

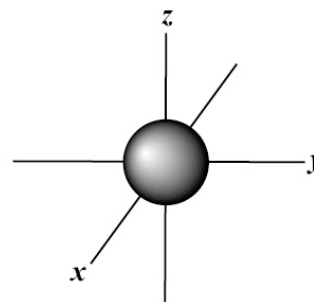
Type of sublevel

$n=1$

1

1s (1 orbital)

2 max e-



Principal
energy
level

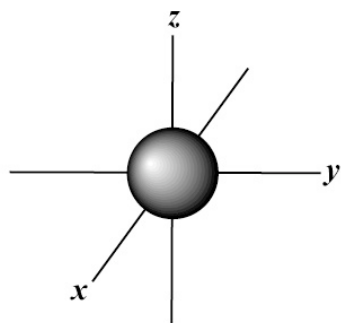
Number
of sublevels

Type of sublevel

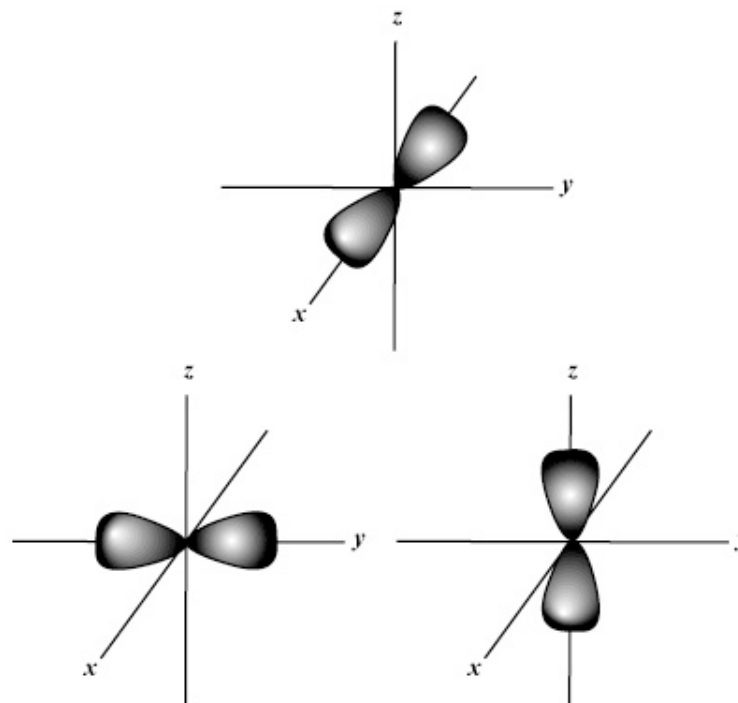
$n=2$

2

2s (1 orbital), 2p (3 orbitals)



8 max e-



Principal
energy
level

$n=3$

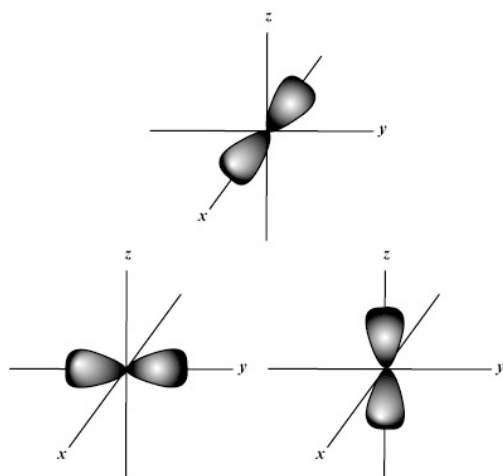
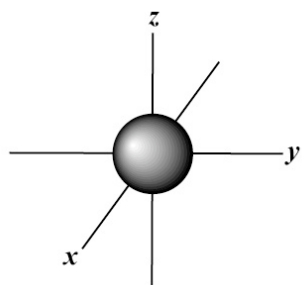
Number
of sublevels

3

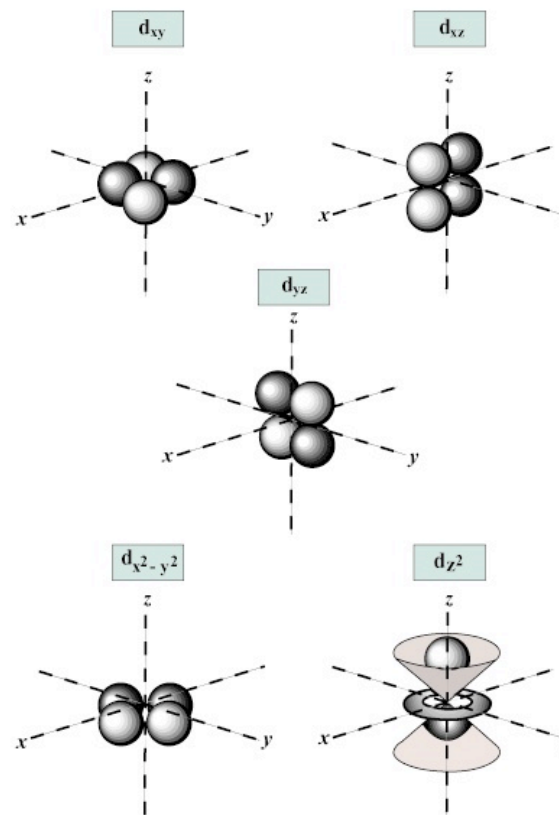
Type of sublevel

3s (1 orbital), 3p (3 orbitals)

3d (5 orbitals)



18 max e-



Principal
energy
level

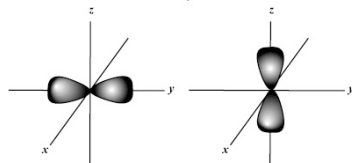
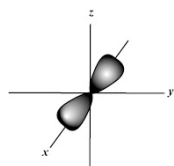
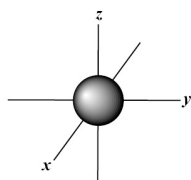
$n=4$

Number
of sublevels

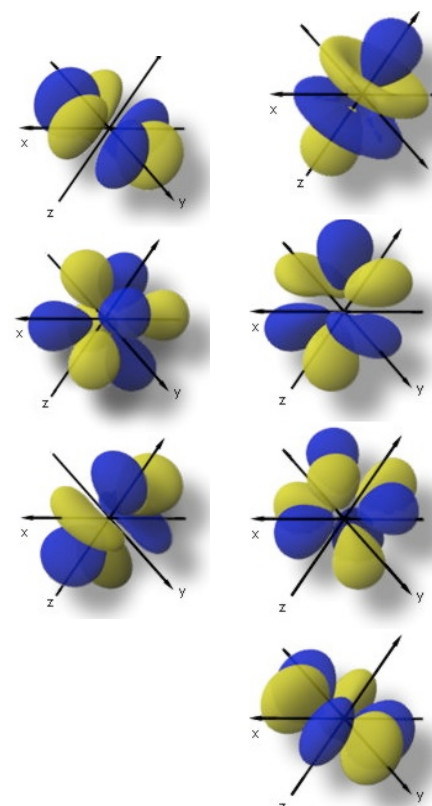
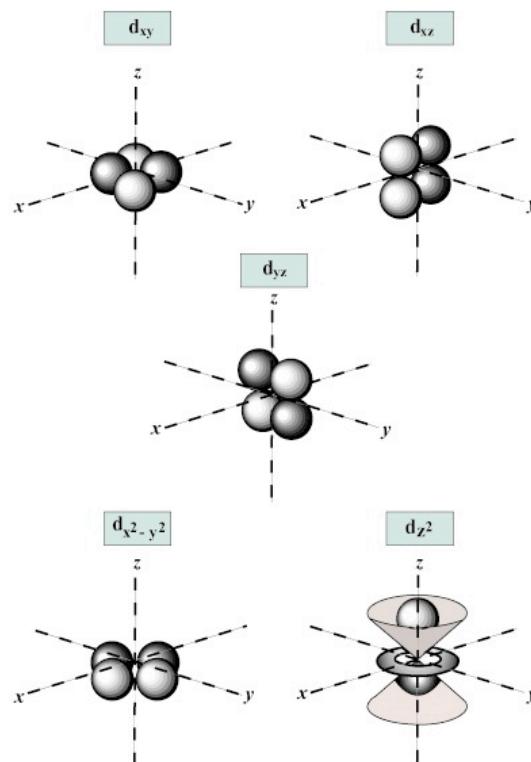
4

Type of sublevel

4s (1 orbital), 4p (3 orbitals)
4d (5 orbitals), 4f (7 orbitals)



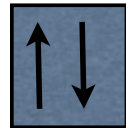
32 max e-



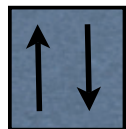
Three rules tell you how to find the electron configuration:

Aufbau principal: e- occupy orbitals of the lowest energy level first.

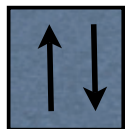
Pauli exclusion principal: an atomic orbital may describe at most 2 e-. Each has a different spin: clockwise and counterclockwise.



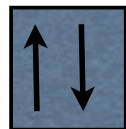
Hund's rule: e- occupy orbitals 1 at a time, with the same spin direction, until all are filled. Only then will e- be paired with another of opposite spin.



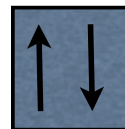
1s



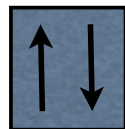
2s



2p_x

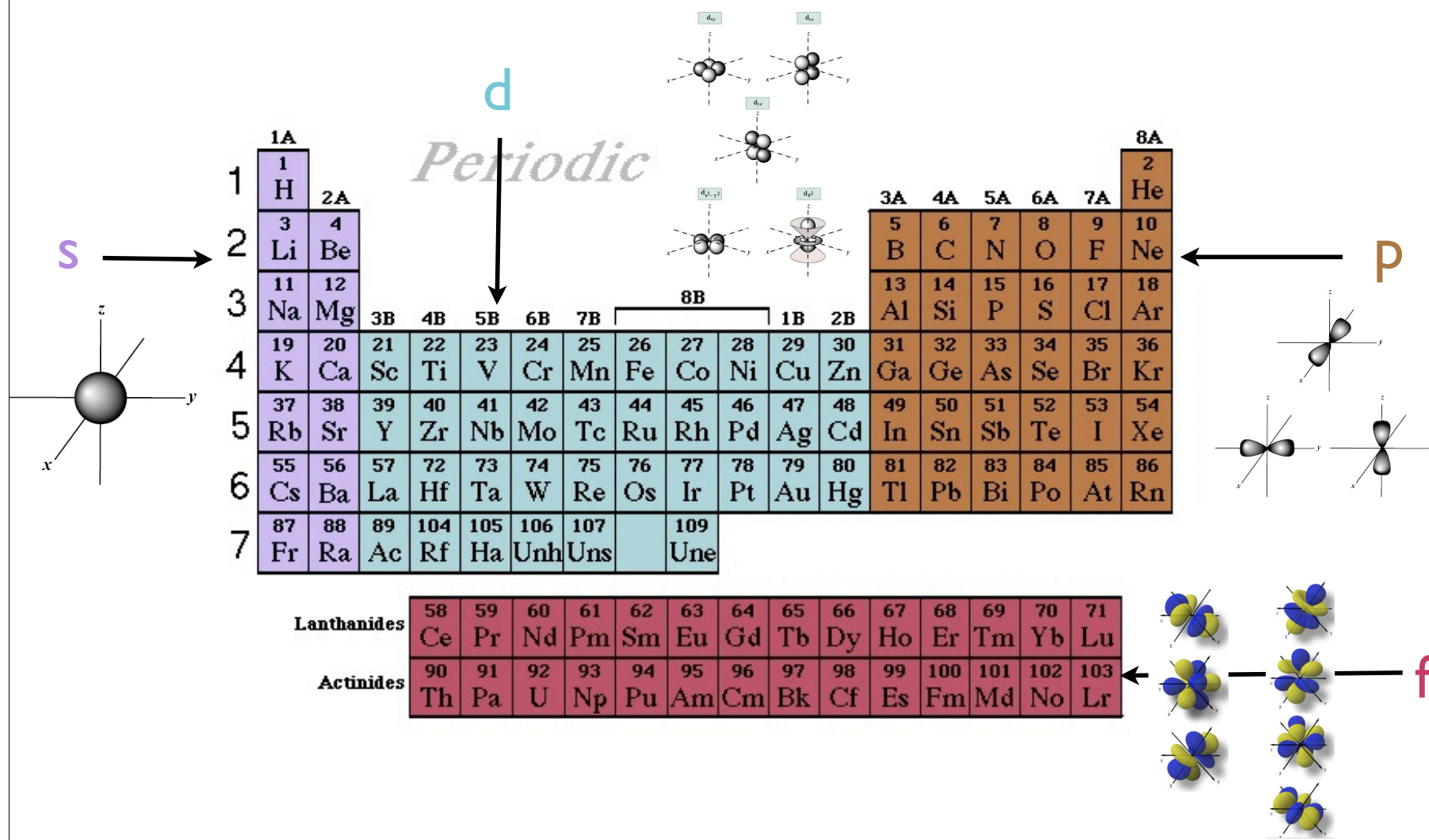


2p_y



2p_z

How do we write electron configurations for each element?



Orbital Filling and Writing Electron Configurations

Periodic Table

1A	1	2A																	8A																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																																														
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Li:



1s



2s



2p_x



2p_y



2p_z

$1s^2 2s^1$

C:



1s



2s



2p_x



2p_y



2p_z

$1s^2 2s^2 2p^2$

Ne:



1s



2s



2p_x



2p_y



2p_z

$1s^2 2s^2 2p^6$

Cl:



1s



2s



2p_x



2p_y



2p_z



3s



3p_x



3p_y



3p_z

$1s^2 2s^2 2p^6 3s^2 3p^5$

Ar:



1s



2s



2p_x



2p_y



2p_z



3s



3p_x



3p_y



3p_z

$1s^2 2s^2 2p^6 3s^2 3p^6$

Orbital Filling and Writing Electron Configurations

Periodic Table

1A	1	H	2A	2	He	8A																													
3	Li	4	Be	5	B	6	C	7	N	8	O	9	F	10	Ne																				
11	Na	12	Mg	13	Al	14	Si	15	P	16	S	17	Cl	18	Ar																				
19	K	20	Ca	21	Sc	22	Ti	23	V	24	Cr	25	Mn	26	Fe	27	Co	28	Ni	29	Cu	30	Zn	31	Ga	32	Ge	33	As	34	Se	35	Br	36	Kr
37	Rb	38	Sr	39	Y	40	Zr	41	Nb	42	Mo	43	Tc	44	Ru	45	Rh	46	Pd	47	Ag	48	Cd	49	In	50	Sn	51	Sb	52	Te	53	I	54	Xe
55	Cs	56	Ba	57	La	58	Ce	59	Pr	60	Nd	61	Pm	62	Sm	63	Eu	64	Gd	65	Tb	66	Dy	67	Ho	68	Er	69	Tm	70	Yb	71	Lu		
87	Fr	88	Ra	89	Ac	90	Th	91	Pa	92	U	93	Np	94	Pu	95	Am	96	Cm	97	Bk	98	Cf	99	Es	100	Fm	101	Md	102	No	103	Lr		

Be:



1s



2s



2p_x



2p_y



2p_z

$1s^2 2s^2$

N:



1s



2s



2p_x



2p_y



2p_z

$1s^2 2s^2 2p^3$

Mg:



1s



2s



2p_x



2p_y



2p_z



$1s^2 2s^2 2p^6 3s^2$

K:



1s



2s



2p_x



2p_y



2p_z



3s



3p_x



3p_y



3p_z



4s

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

Mn:



1s



2s



2p_x



2p_y



2p_z



3s



3p_x



3p_y



3p_z



4s



3d_{xy}



3d_{xz}



3d_{yz}



3d_{x²-y²}



3d_{z²}

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^2$

Let's review electron configurations and write them:

Copper and chromium have configurations that are exceptions to the Aufbau Principle:

Expected configuration:

Cr: $1s^2 2s^2 2p^6 3s^2 3p^6 \underline{3d^4 4s^2}$

Cu: $1s^2 2s^2 2p^6 3s^2 3p^6 \underline{3d^9 4s^2}$

Actual configuration:

$1s^2 2s^2 2p^6 3s^2 3p^6 \underline{3d^5 4s^1}$

$1s^2 2s^2 2p^6 3s^2 3p^6 \underline{3d^{10} 4s^1}$



Having half-filled or filled 3d orbitals is more stable than a filled 4s orbital.

So what is quantum mechanics, really?

It describes the motions of electrons and other subatomic particles as waves.

Electrons move as waves? Aren't they particles???

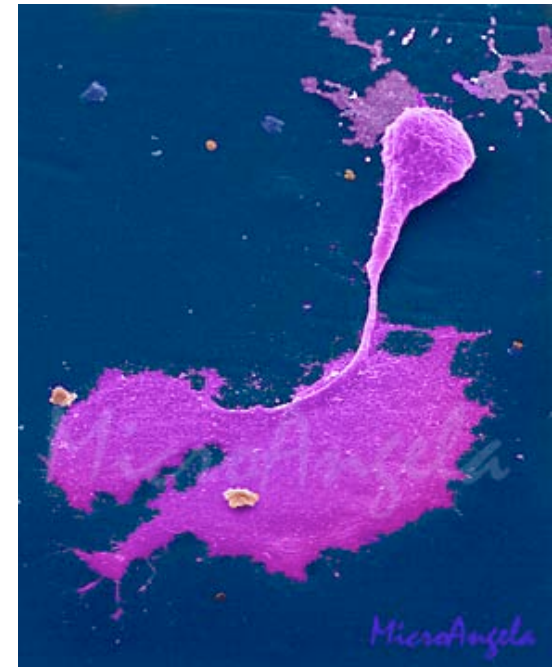
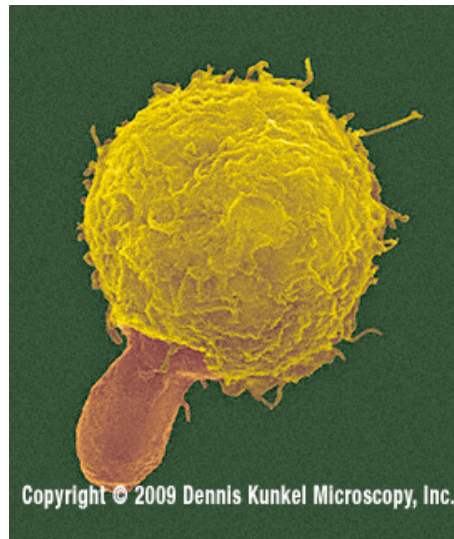
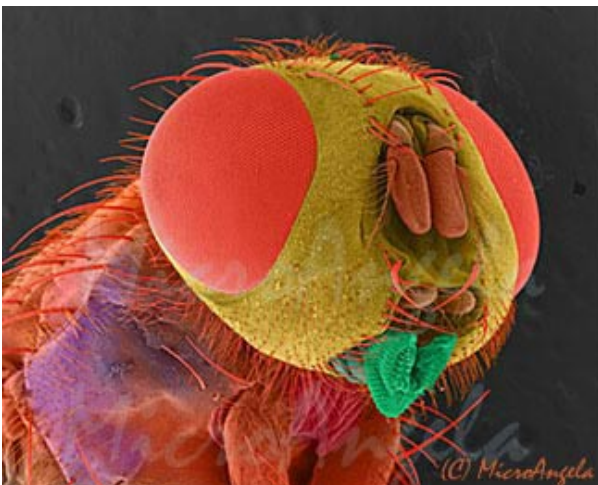
Davisson and Germer, at Bell Labs in NJ, did an experiment that showed that electrons behave as waves!

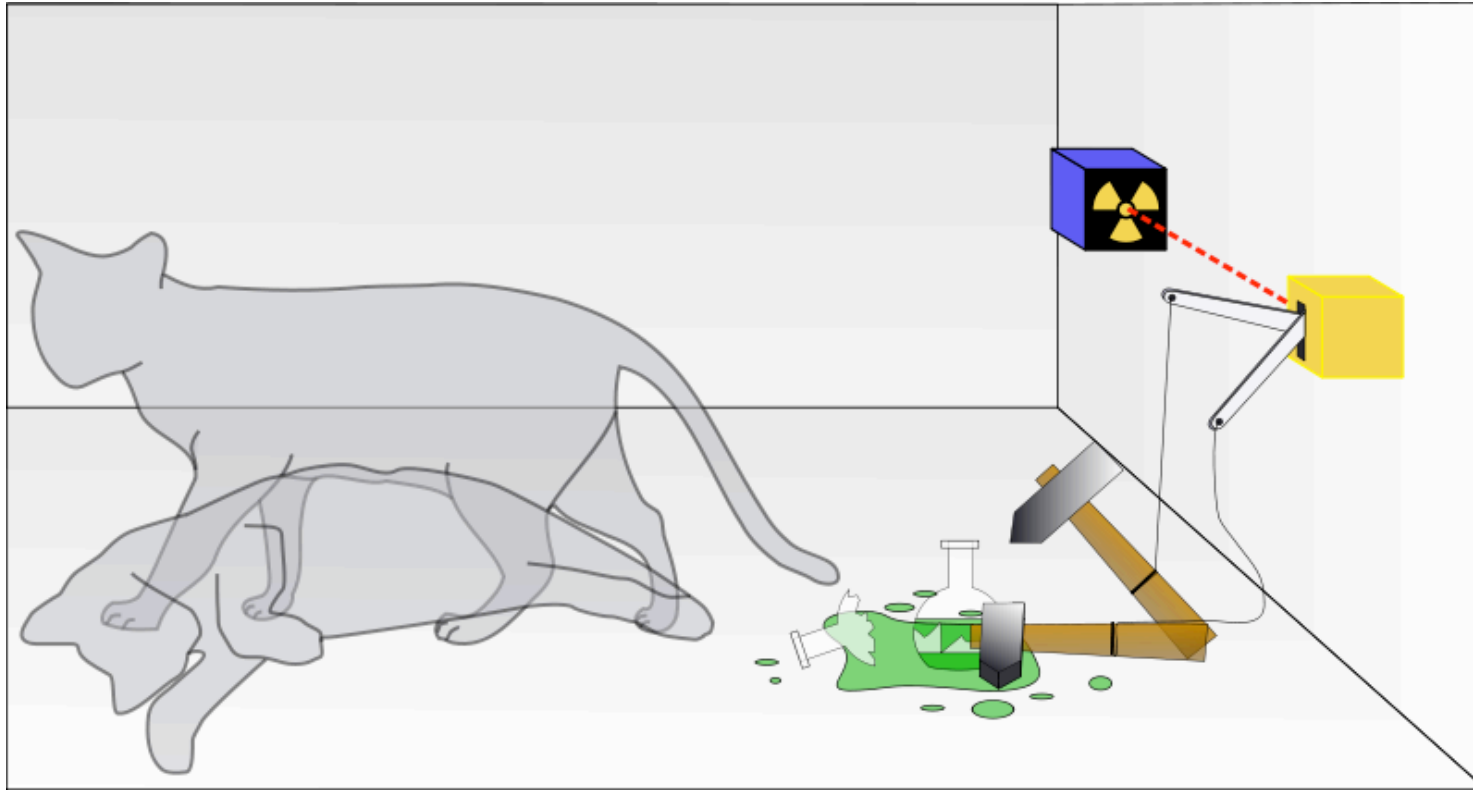
Since electrons behave as waves, they can be used to magnify objects, much like light can....only much smaller objects!

electron microscope



Electron microscope images:





Schrödinger's Cat: A cat, along with a flask containing a poison, is placed in a sealed box shielded against environmentally induced [quantum decoherence](#). If an internal [Geiger counter](#) detects radiation then the flask is shattered, releasing the poison which kills the cat. Quantum mechanics suggests that after a while the cat is [simultaneously alive and dead](#). Yet, when we look in the box, we see the cat *either* alive *or* dead, not a mixture of alive *and* dead.